

Name _____ Period _____ Article # _____

Extra Credit: Gas Laws & SCUBA Diving

This assignment is worth 5 bonus points on your *Gases Test*.

Read the accompanying article "Gas Laws & SCUBA Diving," *ChemMatters*, February 1983, pp. 4-6.

Answer the following questions completely.

1. Why does diving 30m below sea level affect our bodies more than being in a building 30m above sea level?
2. What parts of a diver's body are most affected by pressure changes?
3. State Boyle's Law.
4. Why don't SCUBA diver's lungs collapse as they descend?
5. What would happen to a diver who does not exhale while surfacing from a 30 m dive? Explain in terms of Boyle's Law.
6. State Henry's Law.
7. What gas is associated with causing bubbles in the blood and other body fluids? _____
8. What is another name for decompression sickness? _____
9. Describe how increased pressure in the chamber relieves symptoms of decompression sickness.
10. What is nitrogen narcosis?
11. Which gas law explains why air contaminants (like trace amounts of CO) are more dangerous when the total air pressure is higher? Explain.
12. Is the relationship between the temperature of water and the solubility of a gas in it a direct or inverse relationship? Explain.
13. Use your answer to #12 to explain why it is dangerous for a diver to take a hot shower after a deep dive.

GAS LAWS &

WHAT HAPPENS IF SCUBA DIVERS HOLD THEIR BREATH WHILE MAKING EMERGENCY ASCENTS TO THE SURFACE FROM DEPTHS OF 30 METERS OR MORE?

WHY SHOULDN'T DIVERS FLY OR TAKE HOT SHOWERS SOON AFTER DEEP DIVES?

IS CONTAMINATED COMPRESSED AIR MORE DANGEROUS TO THE DIVER AT THE SURFACE OR AT A DEPTH OF 30 METERS?

Pressure

We live in a sea of air. Since air molecules constantly bombard us, we always experience a pressure of about 760 mm of mercury (or one atmosphere) at the Earth's surface. This is equivalent to 14.7 lb on each square inch of surface. If we zoom to the top of a tall building in an elevator we are no longer as deep in the sea of air as at ground level and, therefore, the pressure around us becomes lower. Ears are usually the first to respond to this change. Wiggling your jaw or swallowing sometimes corrects any discomfort or strange sensations in the ear by opening the tubes connecting the inner ear and throat, allowing the inside pressure to equalize with the outside. A reverse pressure effect is obvious during a rapid airplane descent or during a drive from a mountain pass to the valley floor below.

Divers are surrounded by water molecules in constant motion that exert pressure on their bodies. When you dive to the bottom of the deep end of a

swimming pool, you feel a great deal of pressure exerted by the water. Because water is much more dense than air, pressure changes are much greater for a given change in depth in water than for the same depth change in air. For example, water exerts over 100 lb of force on the surface of a one-gallon metal can pushed just one foot below the water surface. If the metal can contains air, it would not have to be pushed very far below the water surface before the can would start to collapse due to water pressure. Can divers be crushed by the pressure of water in the same manner as the can if they go too deep? After all, for every 10 meters (about 33 ft) in depth, divers experience an additional pressure of one atmosphere.

Pressure-Volume Effects

The changes in pressure experienced by divers are most noticeable on body cavities that contain air, such as the lungs, the middle ear, and the sinus cavities. Boyle's law describes how these gas volumes respond to changes in pressure. For a constant amount of gas at a constant tempera-



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ture, Boyle's law states: *The volume of a gas sample varies inversely with its pressure.*

If divers descend without scuba gear, the amount of gas contained in their body cavities is constant and the volume of these cavities decreases as the surrounding water pressure becomes greater. However, this crushing effect or squeeze is not experienced by divers using scuba gear because the regulator on their air tanks delivers air at the same pressure as the surroundings. This means that the air in divers' lungs is at a pressure equivalent to four atmospheres at a depth of 30 meters. If divers must make emergency ascents from this depth they must remember to breathe out regularly as they return to the surface. If they don't, the pressure of the air in their lungs will cause their lungs to expand. The extreme distortion of the

lungs can cause some of the alveoli (the small sacks in the lungs) to rupture. If this happens, air can enter the bloodstream and cause a blockage that may lead to a variety of problems including loss of consciousness, brain damage, and heart attacks.

The rate of lung expansion increases dramatically as the divers ascend. According to Boyle's law the volume of a flexible gas container will approximately double when the surrounding pressure decreases to one-half its original value. If the divers ascend while holding their breath from a depth of 30 meters (where the pressure is about four atmospheres), their lungs would have to double in volume when they are at 10 meters (where the pressure is about two atmospheres) to equalize the pressure of the water. Of course, this does not happen because the lungs are contained by the rib cage and the muscle system, and the divers are forced to breathe out.

BOYLE'S LAW
 $V \propto \frac{1}{P}$ or $PV = k$

surface = 1 atm

4
units of
volume



10 m = 2 atm

2
units of
volume



20 m = 3 atm

1 1/3
units



30 m = 4 atm

1
unit





Scuba diver examines a piece of the coral on a Caribbean reef.

Pressure-Solubility Effects

Not only does the pressure affect the volume of trapped gases, it also influences the solubility of gases in liquids. Divers must be aware of the principles described by Henry's law, which states: *The amount of gas that will dissolve in a liquid at a given temperature varies directly with the pressure above the liquid.*

Henry's law is useful, therefore, in explaining why during a dive any gases entering the lungs are absorbed to a greater extent in the diver's blood. Although this increased solubility of gases in the blood may create no problems during the dive, the diver's body experiences an effect similar to opening a can of soda when the diver ascends rapidly to the surface. This effect can be accentuated if the diver takes a high-altitude plane flight soon after a dive. In particular, nitrogen gas bubbles that form in the blood and other body fluids can produce a multitude of problems. These problems depend on the location of the gas bubbles, the size and number formed, and the way they are transported by the diver's circulatory system. The bubbles can cause localized pain, itching of the skin, breathing difficulty, and can lead to paralysis, unconsciousness, and death.

To minimize gas bubble formation (decompression sickness or "the bends"), divers carefully follow tables prepared by the U.S. Navy that describe the time limits for dives at various depths greater than 10 meters.

The essence of the process described by the tables involves ascending to a certain point and then remaining at that depth for a time period to allow some of the dissolved nitrogen to escape. Depending on the initial depth, there may be several of these "hold points" during the ascent. If divers experience decompression sickness, the only mode of treatment is to put them in a decompression chamber, increase the pressure surrounding their bodies, and slowly decompress them back to one atmosphere of pressure.

The increased solubility of nitrogen gas at higher pressures may also have a narcotic effect. Nitrogen narcosis or "rapture of the deep" generally does not occur until divers reach depths of about 30 meters. The symptoms are similar in nature to intoxication by alcohol. The divers have a feeling of happiness, overconfidence, tingling or numbness in their arms or legs, and memory impairment. This narcotic effect of nitrogen is just one of the many reasons divers should never work alone underwater.

Another application of Henry's law involves contaminants such as carbon monoxide (CO) that might be present in the compressed air used by divers. Of course, every attempt is made to ensure the purity of the air in scuba tanks, but if a contaminant is present to the extent of just 1%, its presence is more serious during a dive. For example, at a depth of 40 meters, the pressure is equivalent to about five atmospheres. Because the regulator de-

livers air at the same pressure as the surroundings, each breath contains five times more contaminant molecules than each breath from that same tank at the surface. This is equivalent to breathing air containing 5% of that contaminant at the surface.

As the pressure increases during a dive, the solubility of oxygen in the blood also increases proportionately. This means that the effects of poisoning by a trace of carbon monoxide contaminant may go unnoticed during a dive since sufficient oxygen is available for normal cellular respiration. However, as divers surface, the solubility of oxygen decreases in their bloodstreams. Because the carbon monoxide-hemoglobin combination is so stable, there may not be a corresponding decrease of carbon monoxide in the blood. If the divers do not have enough hemoglobin available to bond with oxygen cell respiration, they may lapse into unconsciousness.

Temperature-Solubility Effects

Gas solubility is also affected by changes in temperature. Have you ever noticed that as a cold glass of water warms to room temperature, air bubbles form, clinging to the inside of the glass surface? These bubbles are composed of air that was dissolved in the cooler water. Can you use this information to explain why it is dangerous for a diver to take a hot shower after a deep dive?

A scuba diver with a good basic understanding of gas behavior will better appreciate what is happening during a dive. If you are a scuba diver, this understanding could save your life!

References

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